Directions: (1) Put your name on PART 1 and your name and signature on PART 2 of the exam where indicated.
(2) Sign the Aggie Code on PART 2 of this exam.
(3) Each multiple choice question is actually 2 questions on your scanning sheet. If you are sure of an answer, put the same answer down for both questions for 5 pts. If you cannot decide between two answers, put your best answer down for the first (odd) question and the other answer down for the second (even) question. If you get the first one correct you'll get 3 pts; if you get the second one correct you'll get 2 pts. If there is an ambiguous multiple choice question, use the last page to explain your answer.
(4) Do NOT write on the envelope.
(5) When finished, put everything in the envelope and wait to be excused. At the table, take everything out of the envelope. You can pick up the multiple choice part with the answers outside my office after the exam.
(6) There are a total of 60 multiple choice questions (30 actual questions) plus 20 pts free response.

PART 1

1&2. Which one of the following combinations cannot produce a buffer solution?
(a) HNO₂ and NaNO₂  
(b) NH₃ and (NH₄)₂SO₄  
(c) HClO₄ and NaClO₄  
(d) NH₃ and NH₄Br  
(e) HCN and NaCN

3&4. The rate law for the chemical reaction, 2ClO₂(aq) + 2OH⁻(aq) → ClO₃⁻(aq) + ClO₂⁻(aq) + H₂O(l)
has been determined experimentally to be: Rate = k[ClO₂]²[OH⁻]
The reaction order with respect to the hydroxide ions is:
(a) 3  
(b) 2  
(c) 1  
(d) k[OH⁻]²  
(e) k[ClO₂]²[OH⁻]

5&6. Which is the correct K_c expression for the equilibrium: 2C(s) + O₂(g) ⇌ 2CO(g)?
(a) K_c = \frac{[CO]^2}{2[C][O₂]}  
(b) K_c = \frac{[O₂]^2[O₂]}{[CO]^2}  
(c) K_c = \frac{[CO]^2}{[C][O₂]}  
(d) K_c = \frac{[CO]^2}{[O₂]}  
(e) K_c = \frac{[O₂]}{[CO]^2}

7&8. Which of the following statements is FALSE?
(a) ΔS is a state function.  
(b) A reaction is spontaneous if ΔS_{universe} decreases.  
(c) Endothermic processes are those that absorb heat.  
(d) At absolute 0 K, the entropy of a pure perfect crystalline substance is zero.  
(e) The system's enthalpy alone does not determine the spontaneity of a reaction.
9&10. The best representation for the reaction whose heat of reaction is equal to the standard molar enthalpy of formation for HNO₃(g) is:
(a) H(g) + N(g) + 3O(g) → HNO₃(g)
(b) HNO₃(g) → ½ H₂(g) + ½ N₂(g) + 3/2 O₂(g)
(c) 2HNO₃(g) → H₂(g) + N₂(g) + 3 O₂(g)
(d) H₂(g) + N₂(g) + 3 O₂(g) → 2HNO₃(g)
(e) ½ H₂(g) + ½ N₂(g) + 3/2 O₂(g) → HNO₃(g)

11&12. Which statement is TRUE about aqueous solutions?
(a) As the value of the van’t Hoff factor of the solute increases, the freezing point of the solution will increase.
(b) A solution can only be made by dissolving a solid into a liquid.
(c) When the molality of the solute doubles, the boiling point of the solution will double.
(d) When a solute dissolves into the solvent, the vapor pressure of the solution will be lower than the vapor pressure of the pure solvent at a particular temperature.
(e) A solution is a heterogeneous mixture in which no settling occurs.

13&14. Which of the following species is the STRONGEST oxidizing agent?
(a) Co²⁺  (b) Zn²⁺  (c) Cr²⁺  (d) Cu  (e) Cr

15&16. Consider the following gaseous phase system at equilibrium:
I₂(g) + Cl₂(g) → 2ICl(g) \[\Delta H = -27 \text{ kJ}\]
Which of the following changes will DECREASE the amount of ICl at equilibrium?
(a) adding N₂ (g)
(b) adding a catalyst
(c) decreasing the volume of the container
(d) adding I₂ (g)
(e) raising the temperature

17&18. Consider the following gas phase reaction: A + 2B → AB₂, which occurs by the following mechanism:

Step 1 A + B → AB \hspace{1cm} \text{slow}
Step 2 AB + B → AB₂ \hspace{1cm} \text{fast}
Overall A + 2B → AB₂

The rate law expression must be Rate = ________.
(a) k[A][B]  (b) k[B]  (c) k[A][B]²  (d) k[B]²  (e) k[A]
19&20. The relationship between $K_{sp}$ and $s$, the molar solubility in mol/L, for PbI$_2$ is:

(a) $K_{sp} = 4s^2$  
(b) $K_{sp} = 4s^3$  
(c) $K_{sp} = s^{1/3}$  
(d) $K_{sp} = s^2$  
(e) $K_{sp} = s^3$

21&22. Consider the following reaction and choose the correct statement:

$$
H_2(g) + Br_2(l) \rightarrow 2 HBr(g) \quad \Delta H^0 = -72.8 \text{ kJ} \\
\Delta S^0 = +114 \text{ J/K}
$$

(a) The reaction becomes spontaneous as the temperature increases.  
(b) The reaction becomes nonspontaneous as the temperature increases.  
(c) The reaction is nonspontaneous at all temperatures.  
(d) The reaction is spontaneous at all temperatures.  
(e) The temperature does not influence the spontaneity of any reaction.

23&24. Which of the following statements is **FALSE** about solubility and miscibility?

(a) Iodine, I$_2$(s), should be more soluble in water than in carbon tetrachloride, CCl$_4$.  
(b) Benzene, C$_6$H$_6$, is miscible in chloroform, CHCl$_3$.  
(c) In the phrase “Like dissolves like,” the term “like” refers to molecules with similar polarities.  
(d) Water and ethanol, CH$_3$CH$_2$OH are miscible.  
(e) KCl is more soluble in water than BaS.

25&26. Which one of the following statements is **TRUE** about the following reaction?

$$
CH_3NH_2 + H_2O \xrightleftharpoons{\text{catalyst}} CH_3NH_3^+ + OH^-
$$

(a) OH$^-$ is the conjugate acid of H$_2$O.  
(b) H$_2$O is the conjugate base of OH$^-$.  
(c) CH$_3$NH$_2$ is the conjugate base of H$_2$O.  
(d) CH$_3$NH$_3^+$ is the conjugate acid of CH$_3$NH$_2$.  
(e) There are no conjugate acid-base pairs.

27&28. Which statement or statements is/are **TRUE** about the purpose of a catalyst?

(1) Catalysts increase the number of collisions between the reactants in the overall reaction.  
(2) A catalyst cannot appear as a reactant in a rate law expression.  
(3) A catalyst lowers the activation energy of both the forward and reverse reaction.

(a) 1 & 2  
(b) 2 & 3  
(c) 1 & 3  
(d) only 1  
(e) only 3
29&30. During the electrolysis of aqueous KBr solution using inert electrodes, bromine liquid is evolved at one electrode and hydrogen gas is evolved at the other electrode. The solution around the electrode at which hydrogen gas is evolved becomes basic as the electrolysis proceeds. Which of the following is FALSE?

(a) The electrode where the hydrogen gas is evolved is positively charged.
(b) Faraday's Law says that the longer the cell runs, the more H₂(g) will be produced.
(c) The electrons flow out of the battery into the negatively charged electrode.
(d) The bromide concentration in the cell will decrease.
(e) The electrode where bromine liquid is evolved is the anode.

31&32. Consider the equilibrium, \(2A \rightleftharpoons A_2\) with \(K_c >> 1\). Which picture represents this system?

(a)  
(b)  
(c)  
(d)  
(e)  

33&34. In this titration, what is being titrated with what?

(a) a strong acid is being titrated with a strong base
(b) a strong base is being titrated with a strong acid
(c) a weak acid is being titrated with a strong base
(d) a weak base is being titrated with a strong acid
(e) none of the above.
35&36. What is the pH of a $1.5 \times 10^{-4}$ M KOH?

(a) 2.95  (b) 3.80  (c) 10.18  (d) 10.79  (e) 11.52

37&38. What is the pH of a 0.70 M formic acid solution?

(a) 4.19  (b) 3.27  (c) 2.32  (d) 4.85  (e) 1.95

39&40. Rate data were collected for the following reaction at a particular temperature. What is the correct rate law expression?

$$A + 2B \rightarrow C + 3D$$

<table>
<thead>
<tr>
<th>Experiment</th>
<th>$[A]_{initial}$</th>
<th>$[B]_{initial}$</th>
<th>Initial Rate of Reaction</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>0.10 M</td>
<td>0.40 M</td>
<td>0.0090 M/min</td>
</tr>
<tr>
<td>2</td>
<td>0.20 M</td>
<td>0.40 M</td>
<td>0.036 M/min</td>
</tr>
<tr>
<td>3</td>
<td>0.10 M</td>
<td>0.20 M</td>
<td>0.0045 M/min</td>
</tr>
</tbody>
</table>

(a) Rate = $k[A][B]$  (b) Rate = $k[A]^2[B]$  (c) Rate = $k[A][B]^2$

(d) Rate = $k[A]^2[B]^2$  (e) Rate = $k[A]^2$
41&42. If a solution is $1.0 \times 10^{-2}$ M in Mg$^{2+}$ and $1.0 \times 10^{-4}$ M in F$^-$, will precipitation occur?

(a) No, because Qsp < Ksp  
(b) No, because Ksp < Qsp  
(c) Yes, because Qsp < Ksp  
(d) Yes, because Ksp < Qsp  
(e) cannot be determined

43&44. Calculate the standard cell potential for the cell: Fe/FeSO$_4$ (1 M) || CuSO$_4$ (1 M) /Cu

(a) +0.65 V  
(b) +0.15 V  
(c) +0.06 V  
(d) +0.32 V  
(e) +0.78 V
45&46. If 4.27 grams of sucrose, \( \text{C}_{12}\text{H}_{22}\text{O}_{11} \) are dissolved in 16.2 grams of water, what will be the boiling point of the resulting solution?

(a) 101.64°C   (b) 100.39°C   (c) 99.626°C   (d) 100.73°C   (e) 101.42°C

47&48. How many grams of chromium metal can be made from a solution of \( \text{Cr(NO}_3\text{)}_3 \) when a current of 4.00 amperes flows for 50.0 minutes?

(a) 3.68 g   (b) 1.43 g   (c) 2.16 g   (d) 0.824 g   (e) 1.63 g
49&50. What is the enthalpy change of the reaction below at 298 K and 1 atm pressure?

\[
\begin{array}{cccc}
\text{SiH}_4(s) & + & 2\text{O}_2(g) & \rightarrow \\
\text{SiO}_2(s) & + & 2\text{H}_2\text{O}(g)
\end{array}
\]

| \(\Delta H_{298}^o (\text{kJ/mol})\) | +34 | 0 | −910.9 | −285.8 |

(a) −1159 kJ  
(b) −955 kJ  
(c) −1230 kJ  
(d) −1250 kJ  
(e) −1517 kJ

51&52. What concentration of \(\text{Pb}^{2+}\) will initiate precipitation in a solution that is 0.010 M NaBr?

(a) 6.3 \times 10^{-1} \text{ M}  
(b) 6.3 \times 10^{-4} \text{ M}  
(c) 6.3 \times 10^{-6} \text{ M}  
(d) 6.3 \times 10^{-2} \text{ M}  
(e) 6.3 \times 10^{-5} \text{ M}
53&54. What ratio of $[\text{HNO}_2]/[\text{NO}_2^-]$ is required to give a solution with a pH of 4.00?

(a) 4.6:1  (b) 0.22:1  (c) 1:1  (d) 2.7:1  (e) 0.15:1

55&56. Calculate the pH of a 0.25 M NaBrO solution.

(a) 11.00  (b) 9.69  (c) 8.50  (d) 9.46  (e) 10.12
57&58. Consider the following equilibrium reaction: \( \text{N}_2(g) + \text{O}_2(g) \rightleftharpoons 2\text{NO}(g) \) \hspace{1cm} \( K_c = 4.0 \times 10^{-4} \) at 2010 K

If the initial concentration of NO in a closed container is 3.00 M, what will be the concentration of NO after the system finally reaches equilibrium at 2010 K?

(a) 0.49 M  
(b) 0.10 M  
(c) 0.007 M  
(d) 0.030 M  
(e) 1.30 M

59&60. Consider the reaction: \( 2 \text{CO}(g) \rightarrow \text{C(s, graphite)} + \text{CO}_2(g) + 172.5 \text{ kJ} \)

At 25°C, the \( \Delta S^\circ \) for the reaction is \( -175.9 \text{ J/K} \). Calculate \( \Delta G^\circ \) for the reaction at 25°C.

(a) \(-178 \text{ kJ}\)  
(b) \(-197 \text{ kJ}\)  
(c) \(-120. \text{ kJ}\)  
(d) 0.981 kJ  
(e) 4620 kJ
PART 2

Please read and sign: “On my honor, as an Aggie, I have neither given nor received unauthorized aid on this exam.” ________________________________

(5 pts) **61.** Calculate the potential (in volts) for the non-standard voltaic cell when the following two half-cells are connected: Anode: Ag electrode in 1.0 x 10^-3 Ag^+ solution  
Cathode: Au electrode in 0.10 M Au^{3+} solution

Given:  \( E_{\text{cell}} = E_{\text{cell}}^0 - (0.0257/n) \ln Q_{\text{thermo}} \)  
\( E_{\text{cell}} = E_{\text{cell}}^0 - (0.0592/n) \log Q_{\text{thermo}} \)

(5 pts) **62.** Assign oxidation numbers to every element in the reaction (2 pts). Balance the following redox reaction in acidic solution (3 pts):

\[ \text{Cu} + \text{NO}_3^- \rightarrow \text{Cu}^{2+} + \text{NO} \]
63. A buffer is prepared by mixing 100 mL of 0.200 M ammonia and 300 mL of 0.100 M ammonium chloride. What is the pH of the final solution after 20.0 mL of 0.400 M HCl is added?

64. Draw the voltaic cell that results when the following 2 half cells are connected:
   (1) an iron electrode is put into a solution of 1.00 M Fe$^{2+}$ solution and
   (2) a lead electrode is put into a solution of 1.00 M Pb$^{2+}$ solution.

Observations:
   (1) The iron electrode decreases in mass while the [Fe$^{2+}$] increases
   (2) The lead electrode increases in mass while the [Pb$^{2+}$] decreases

(1 pt) Label the anode and give the anodic reaction.
(1 pt) Label the cathode and give the cathodic reaction?
(1 pt) Give the sign on each electrode?
(1 pt) Show the direction of the electron flow.
(1 pt) What is the overall reaction?